Date: ____________________ Name: _______________________

LABELING ELECTRONS IN ATOMS

The location of each electron in an atom is determined by a few different factors. Each factor is represented by a QUANTUM NUMBER.

Prediction: What do you think the different factors might be?

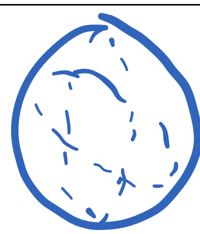
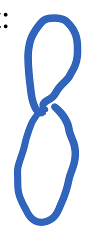
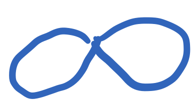
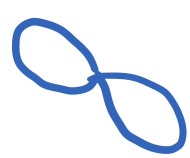
- energy level
- orbital shape

FACTOR ONE: Energy Level; relative size of the atomic orbital
→ PRINCIPAL QUANTUM NUMBER, n
→ $n = 1, 2, 3, 4, \ldots$

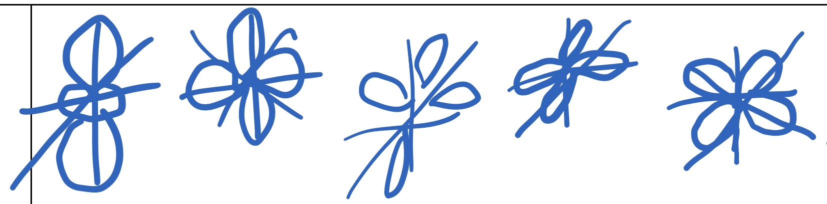
The higher the value of n, the higher (higher OR lower) the energy level

FACTOR TWO: Orbital Shape
→ Angular momentum quantum number, ℓ
→ ℓ = s, p, d, f

FACTOR THREE: orientation in space
→ Magnetic quantum number (m_ℓ)
→ x, y or z (ex. $2p_x$, $2p_y$, $2p_z$)

Letter	Orbital Shape		
s	Draw it:  (sphere)		
p	Draw it: 	Draw it: 	Draw it: 

Date: ____________________ Name: ________________________

d	
f	

Linking Ideas: Write down a pneumonic that will help you remember the order of the letters representing the different orbital shapes. So People Don't Forget

So, how do we know the energy levels and orbital shapes of all the electrons in an atom?

First, let's summarize what we know: (Fill in any blanks.)

Energy Level	**Sublevels**	**Number of Orbitals**	**Number of Electrons**	**Total Number of Electrons**
1	s	1	2	2
2	s	1	2	8
	p	3	6	
3	s	1	2	18
	p	3	6	
	d	5	10	
4	s	1	2	32
	p	3	6	
	d	5	10	
	f	7	14	

<u>FACTOR FOUR:</u> electron spins ↑↓ → electrons

→ Spin quantum number (m_s); Spin is either + ½ or – ½

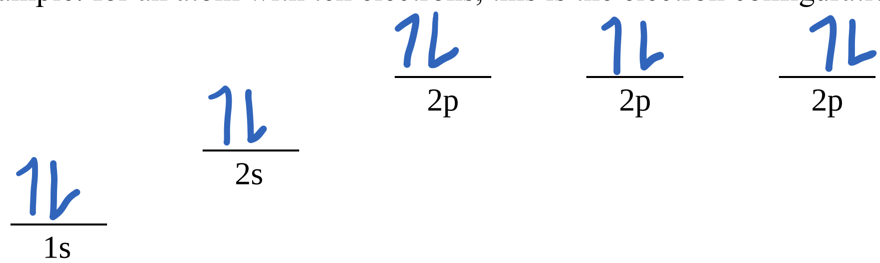
Spins in atoms result in analysing OPPOSING magnetic fields.

Date: ____________________ Name: _______________________

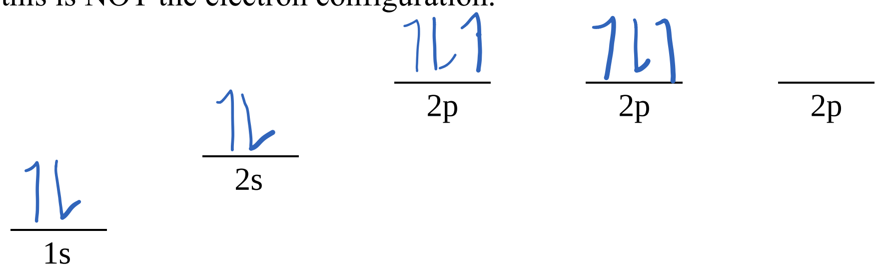
To find out the electron positions and orientations in different atoms, we just need a few rules.

1. PAULI EXCLUSION PRINCIPLE:
   Only two electrons can fit into each orbital.

   For example: for an atom with ten electrons, this is the electron configuration…

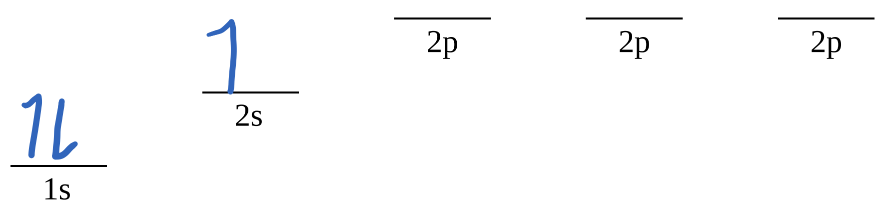


   …and this is NOT the electron configuration.

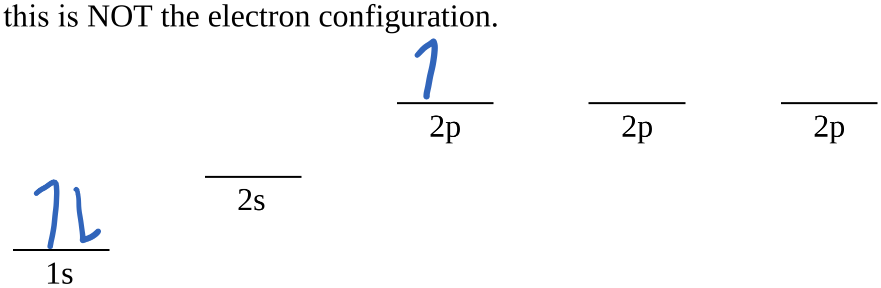


2. AUFBAU PRINCIPLE:
   An electron will always occupy an orbital of the LOWEST possible energy.

   For example: for an atom with three electrons, this is the electron configuration…



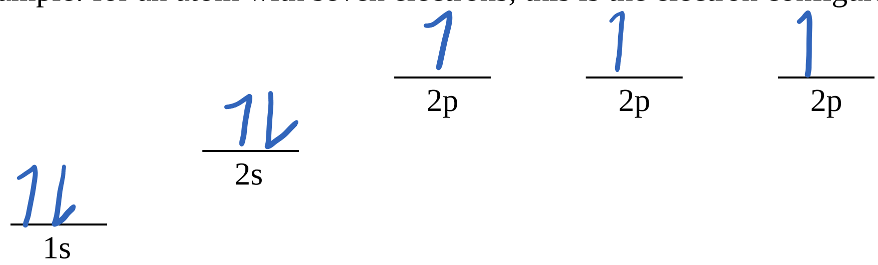
   …and this is NOT the electron configuration.



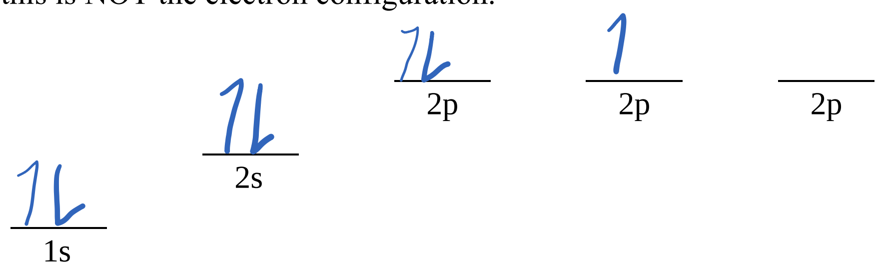
Date: ____________________ Name: _______________________

3. HUND'S RULE:
All orbitals of equal energy must be occupied by one electron before any pairing occurs in those orbitals.

For example: for an atom with seven electrons, this is the electron configuration...

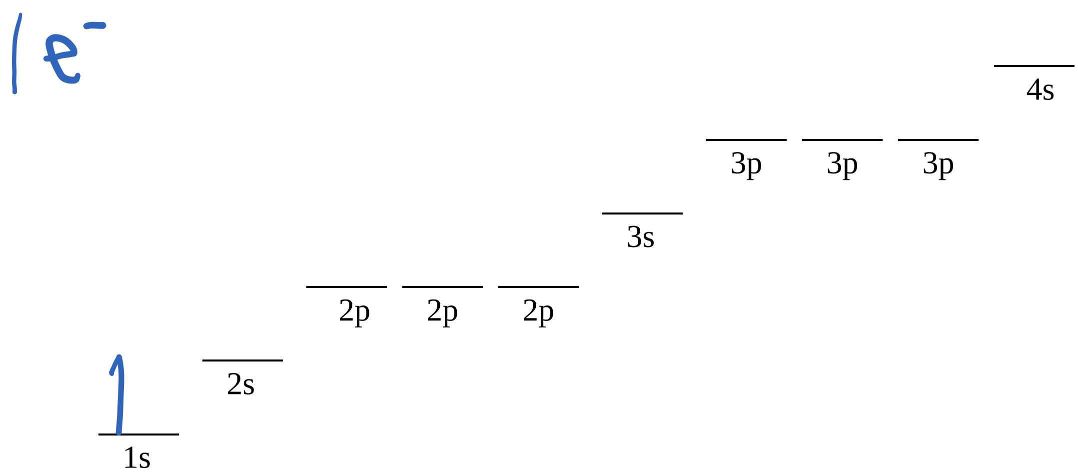


…and this is NOT the electron configuration.



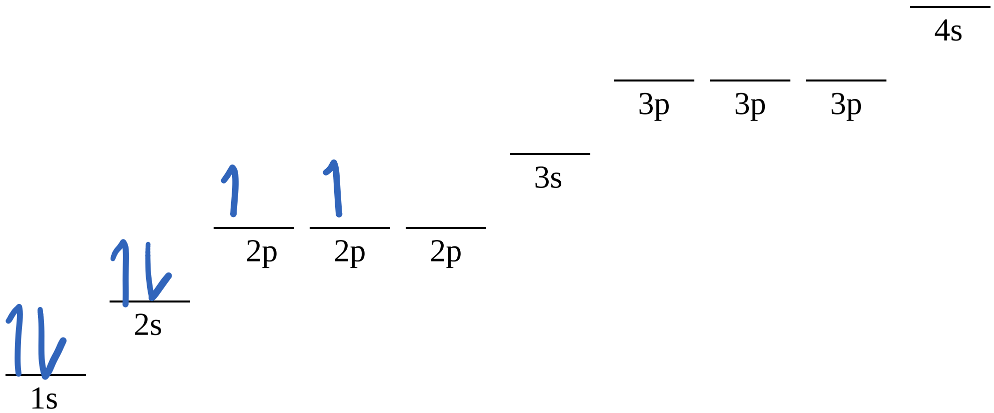
Application: Use the following diagrams to draw orbital diagrams for these atoms: H, C, Na and S.

H:

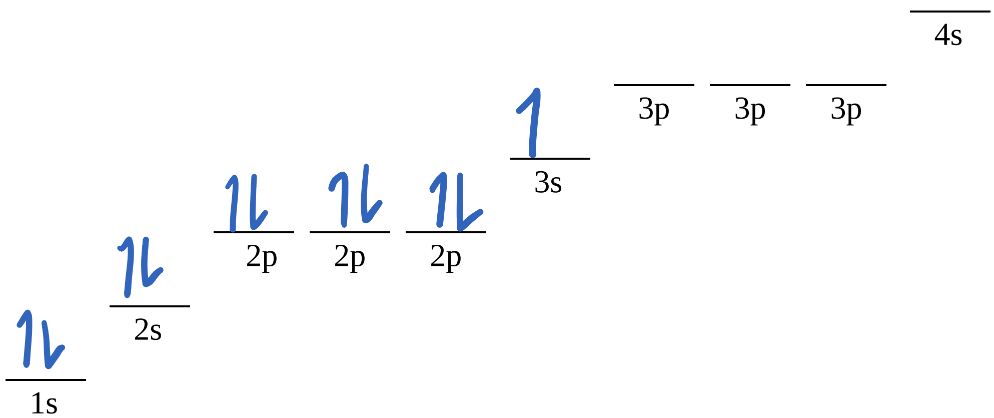


Date: ____________________   Name: ________________________

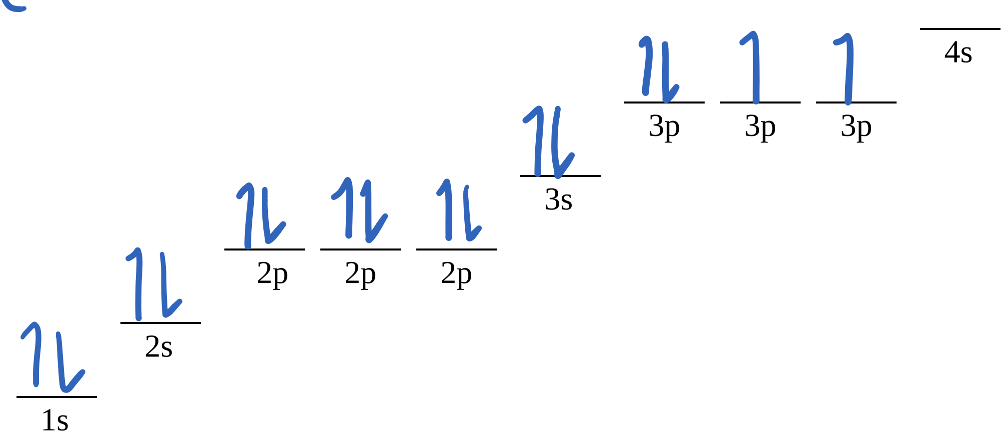
C: 6 e^-



Na: 11 e^-



S: 16 e^-



Date: ____________________ Name: _______________________

These orbital diagrams can be summarize as electron configurations as follows:

H: $1s^1$

C: $1s^2\,2s^2\,2p^2$

Na: $1s^2 2s^2 2p^6 3s^1$

S: $1s^2 2s^2 2p^6 3s^2 3p^4$

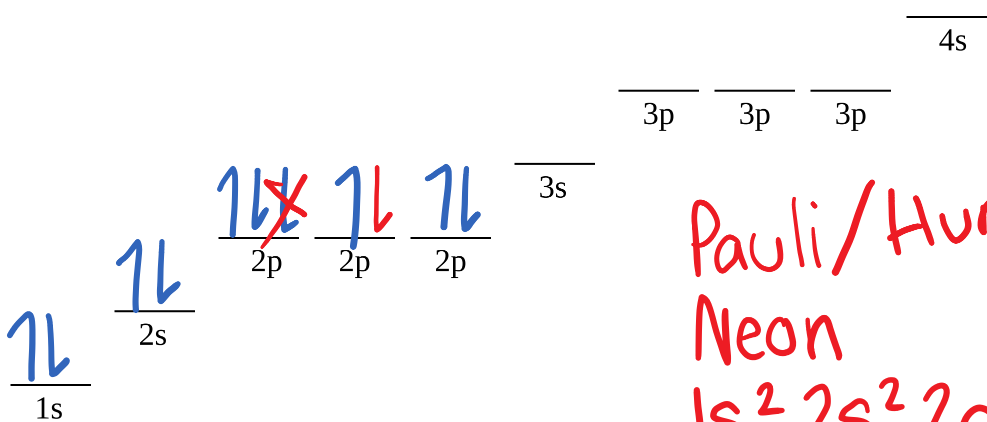
# of electrons in orbital ↓

energy level → $1s^2$ ← orbital shape

Expanding Ideas: For each of the following orbital diagrams

a) Determine which rule is being broken.
b) Correct the orbital diagram (use a coloured pen!)
c) Write the correct electron configuration.
d) Determine which element it is.

1.

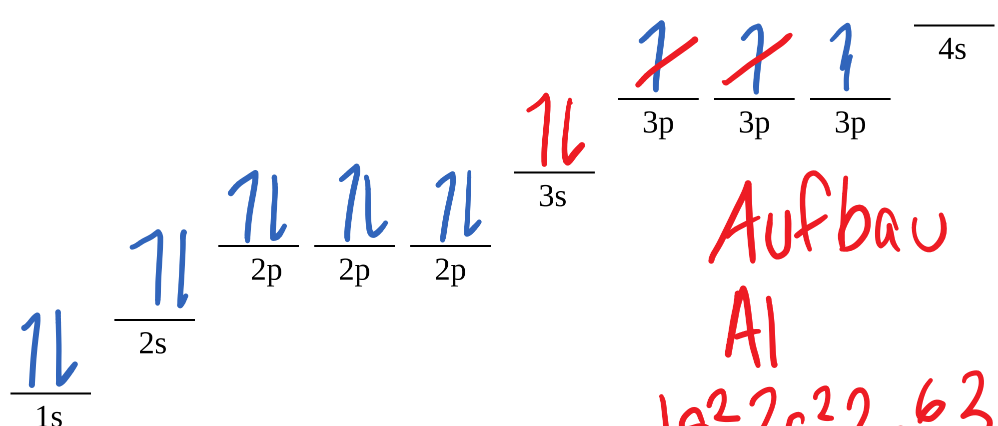


Pauli/Hunds

Neon

$1s^2 2s^2 2p^6$

2.



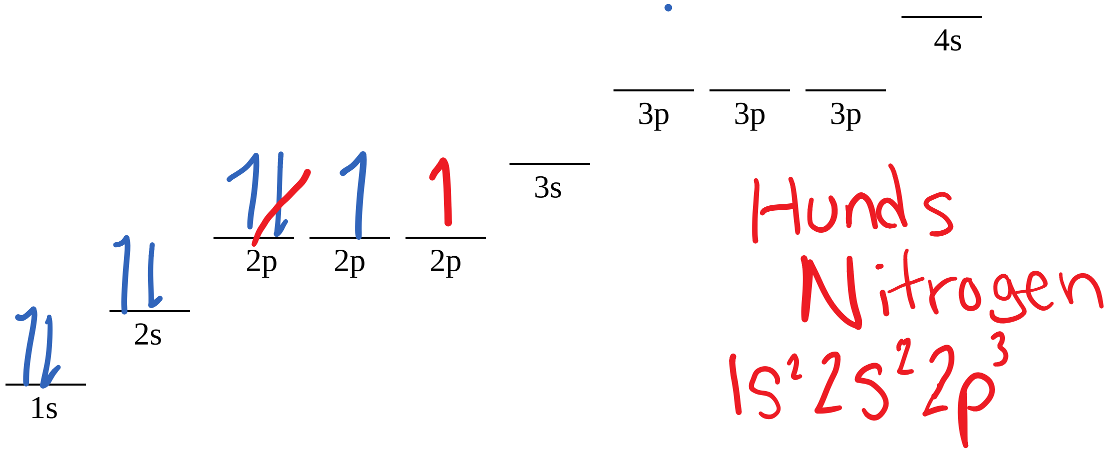
Aufbau

Al

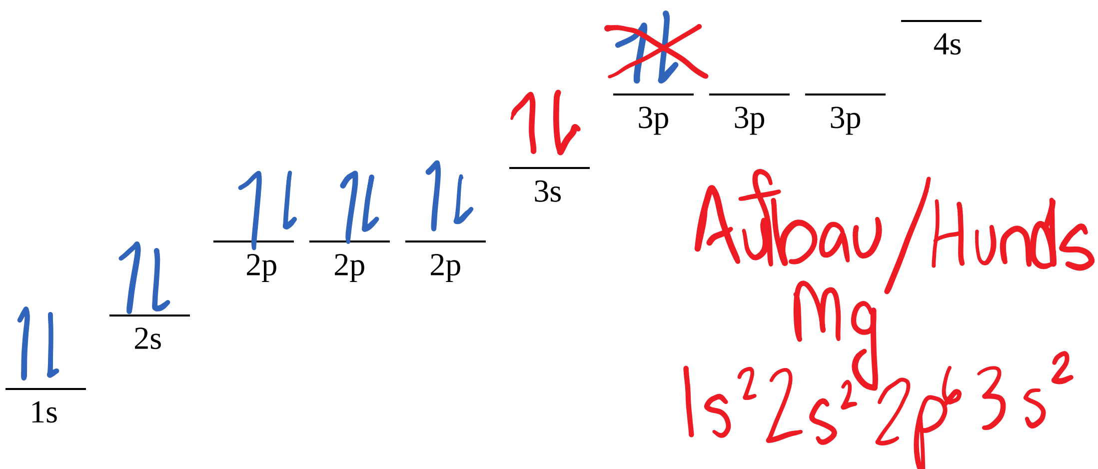
$1s^2 2s^2 2p^6 3s^2 3p^1$

Date: ____________________ Name: ______________________

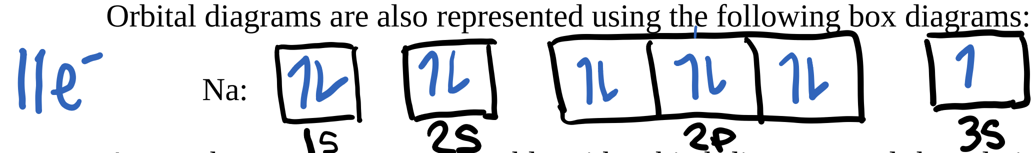
3.



4.



Orbital diagrams are also represented using the following box diagrams:



As you become more comfortable with orbital diagrams and the relative energies of the different orbitals, you may want to condense your work using the above format.

Practice: Write electron configurations for He, Li, Ne and Na.

He : $1s^2$

Li: $1s^2 2s^1$

Ne: $1s^2 2s^2 2p^6$

Na: $1s^2 2s^2 2p^6 3s^1$

Date: ____________________ Name: _______________________

Linking Ideas: What do you notice about the electron configurations of

a) He and Li?

same e^- configuration but Li has one more e^-

b) Ne and Na?

" " but Na has one more e^-

NOTE: Electron configurations can be written in terms of the noble gases (He, Ne, Ar, Kr,…)

Li: [He] $2s^1$

Na: $[Ne]\,3s^1$

Practice: Write the electron configurations, in terms of noble gases, for Mg, Ca and Be.

Mg: $[Ne]\,3s^2$

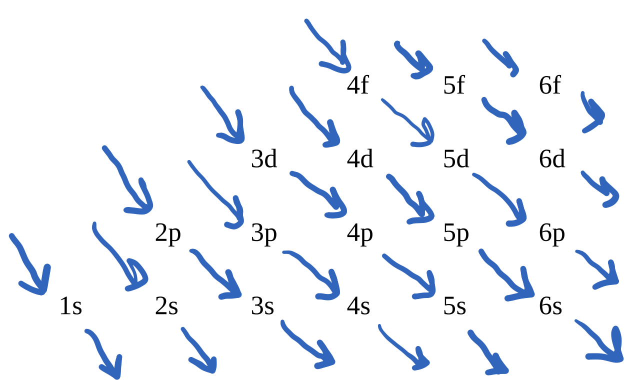
Ca: $[Ar]\,4s^2$

Be: $[He]\,2s^2$

As the energy levels get higher, there is a change in the pattern of orbitals:

4d
5s
4p
3d
4s
3p
3s
2p
2s
1s

This can be remembered using this diagram:



Practice: Write electron configurations for Li, F, Na, Cl, K and Br.

Li: $1s^2 2p^1$

F: $1s^2 2s^2 2p^5$

Na

Cl: $1s^2 2s^2 2p^6 3s^2 3p^5$

K: $[Ar]\,4s^1$

Br: $[Ar]\,4s^2 3d^{10} 4p^5$

Date: ____________________ Name: _______________________

*There are two important electron configuration exceptions (Cr and Cu) that you are responsible for knowing and explaining.

Cr: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$
Cu: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$
} half & full d shells are more stabe

Extension: To write electron configurations for ions, following these guidelines:

a) negative ions: add electrons to the last unfilled subshell, starting where the neutral atom left off.

F^-: $1s^2 2s^2 2p^{\not 5\, 6}$

18e⁻ S^{2-}: $1s^2 2s^2 2p^6 3s^2 3p^{\not 4\, 6}$

b) positive ions: i) remove electrons in the outermost shell (largest *n*-value) first.
ii) within a given outermost shell, remove electrons in the order: *p*–electrons, *s*–electrons, *d*–electrons

Sn^{2+}: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 | 5s^2 4d^{10} \not{5p^2}$

Sn^{4+}: [Kr] $4d^{10}$

Nb^{2+}: [Kr] $4d^3$